T05D07 – IB SL Question Review

Name KEY

1. Consider the specific heat capacity of the following metals.

Metal	Specific heat capacity / J kg^{-1} K^{-1}
Cu	385
Ag	234
Au	130
Pt	134

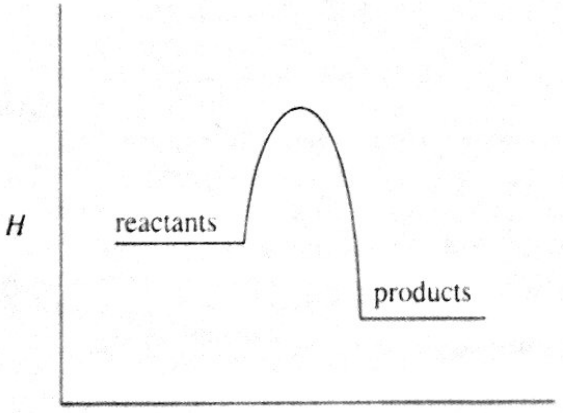
Which metal will show the greatest temperature increase if 50 J of heat is supplied to a 0.001 kg sample of each metal at the same initial temperature?

A. Cu
B. Ag
C. Au
D. Pt

2. Which statement is correct for an endothermic reaction?

A. Bonds in the products are stronger than the bonds in the reactants.
B. Bonds in the reactants are stronger than the bonds in the products.
C. The enthalpy of the products is less than that of the reactants.
D. The reaction is spontaneous at low temperatures but becomes non-spontaneous at high temperatures.

3. According to the enthalpy level diagram below, what is the sign for ΔH and what term is used to refer to the reaction?



	ΔH	reaction
A.	positive	endothermic
B.	negative	exothermic
C.	positive	exothermic
D.	negative	endothermic

4. When 40 joules of heat are added to a sample of solid H_2O at −16.0°C the temperature increases to −8.0°C. What is the mass of the solid H_2O sample?

[Specific heat capacity of $H_2O(s) = 2.0\ J\ g^{-1}K^{-1}$]

A. 2.5 g
B. 5.0 g
C. 10 g
D. 160 g

5. The $\Delta H^{\ominus}$ values for the formation of two oxides of nitrogen are given below.

$$\tfrac{1}{2}N_2(g) + O_2(g) \rightarrow NO_2(g) \qquad \Delta H^{\ominus} = -57\ kJ\ mol^{-1}$$

$$N_2(g) + 2O_2(g) \rightarrow N_2O_4(g) \qquad \Delta H^{\ominus} = +9\ kJ\ mol^{-1}$$

Use these values to calculate $\Delta H^{\ominus}$ for the following reaction (in kJ):

$$2NO_2(g) \rightarrow N_2O_4(g)$$

A. −105
B. −48
C. +66
D. +123

6. How much energy, in joules, is required to increase the temperature of 2.0 g of aluminium from 25 to 30°C?

(Specific heat of Al = 0.90 J g^{-1} K^{-1})

A 0.36
B 4.5
C 9.0
D 54

(Total 1 mark)

7. Which combination is correct for a chemical reaction that absorbs heat from the surroundings?

	Type of reaction	ΔH at constant pressure
A	Exothermic	Positive
B	Exothermic	Negative
C	Endothermic	Positive
D	Endothermic	Negative

(Total 1 mark)

8. Using the equations below:

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H^{\ominus} = -394$ kJ mol^{-1}

$Mn(s) + O_2(g) \rightarrow MnO_2(s)$ $\Delta H^{\ominus} = -520$ kJ mol^{-1}

What is ΔH, in kJ, for the following reaction?

$MnO_2(s) + C(s) \rightarrow Mn(s) + CO_2(g)$

A 914
B 126
C −126
D −914

(Total 1 mark)

9. (a) Define the term *average bond enthalpy*, illustrating your answer with an equation for methane, CH_4.

Energy, for the conversion of a gaseous molecule into (gaseous) atoms, the average value is obtained by mass spectrometry from a number of different bonds/compounds

$CH_4 \rightarrow C(g) + 4H(g)$ *Need state symbols!*

(3)

(b) The equation for the reaction between methane and chlorine is

$CH_4(g) + Cl_2(g) \rightarrow CH_3Cl(g) + HCl(g)$

Use the values from Table 10 of the Data Booklet to calculate the enthalpy change for this reaction.

H − C − H + Cl − Cl → H − C − Cl + H − Cl

(C−H) × 4 (413)
(Cl−Cl) × 1 (242)

Σ(reactants) − Σ(products)

(C−H) × 3 (413)
(C−Cl) × 1 (328)
(H−Cl) × 1 (432)

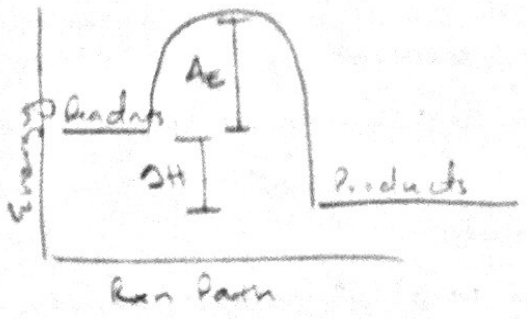
(3)

(c) Explain why no reaction takes place between methane and chlorine at room temperature unless the reactants are sparked, exposed to UV light or heated.

Molecules have insufficient energy to react (at room T)
(98%) wrong collision geometry (ineffective) unsuccessful collisions
(2%) enough energy, is needed to overcome the activation energy (E_a) for the reaction

(2)

(d) Draw an enthalpy level diagram for this reaction.



(2)
(Total 10 marks)

10. In aqueous solution, potassium hydroxide and hydrochloric acid react as follows.

$KOH(aq) + HCl(aq) \rightarrow KCl(aq) + H_2O(l)$

The data below is from an experiment to determine the enthalpy change of this reaction.

50.0 cm^3 of a 0.500 mol dm^{-3} solution of KOH was mixed rapidly in a glass beaker with 50.0 cm^3 of a 0.500 mol dm^{-3} solution of HCl.

Initial temperature of each solution = 19.6°C

Final temperature of the mixture = 23.1°C

(a) State, with a reason, whether the reaction is exothermic or endothermic.

exothermic b/c temperature rises
(OR) heat is released

(1)

(b) Explain why the solutions were mixed rapidly.

To make any heat loss as small as possible (OR) so that all the heat will be given out very quickly

NOT → "to produce a faster reaction" ← NOT ACCEPTABLE!

(1)

(c) Calculate the enthalpy change of this reaction in kJ mol^{-1}. Assume that the specific heat capacity of the solution is the same as that of water.

heat released = mass × specific heat × temperature increase.

$q_{H_2O} = mc\Delta T$

$q_{H_2O} = (100g)(4.184\ J/g^{\circ}C)(3.5^{\circ}C)$

$q_{H_2O} = 1463\ J$ or $1.463\ kJ$

$q_{HCl} = -q_{H_2O}$

$q_{KOH} = -1.463\ kJ$

$\frac{-1.463\ kJ}{0.025\ mol} = \boxed{-58.5\ kJ/mol}$

To find mol:

0.500 M HCl = $\frac{x\ mol}{0.0500\ L}$

0.500 M KOH = $\frac{x\ mol}{0.0500\ L}$

in both cases mol = 0.025 and so are w/ respect to q, either

(4)

mass of H_2O:
assume dissolved ions do NOT affect specific heat or density of water. So,
KOH(aq) = 50mL
HCl(aq) = 50mL
H_2O = 100mL = 100g

(d) Identify the **major** source of error in the experimental procedure described above. Explain how it could be minimized.

Heat loss to the surroundings

Use insulation on the reaction vessel (OR) use a lid (OR) draw temperature vs. time graph.

(2)

(e) The experiment was repeated but with an HCl concentration of 0.510 mol dm^{-3} instead of 0.500 mol dm^{-3}. State and explain what the temperature change would be.

Still 3.5°C = ΔT

B/c KOH is still the limiting reactant and the excess HCl would not react.

(2)
(Total 10 marks)

11. Given the following data:

× 1 $C(s) + 2F_2(g) \rightarrow CF_4(g)$, $\Delta H_1 = -680\ kJ\ mol^{-1}$

× −2 $F_2(g) \rightarrow 2F(g)$, $\Delta H_2 = +158\ kJ\ mol^{-1}$

× −1 $C(s) \rightarrow C(g)$, $\Delta H_3 = +715\ kJ\ mol^{-1}$

calculate the average bond enthalpy (in kJ mol^{-1}) for the **C—F** bond.

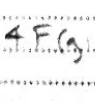
$C(s) + 2F_2(g) \rightarrow CF_4(g)$ (−680)

$4F(g) \rightarrow 2F_2(g)$ (+158) × −2

$C(g) \rightarrow C(s)$ (+715) × −1

$4F(g) + C(g) \rightarrow CF_4(g)$ $\Delta H_{rxn} = -1711\ kJ$ / 4 C–F bonds

(Total 4 marks)



$\boxed{\Delta H_{C-F} = -428\ kJ/mol}$

* Ok to answer w/ +/− value in range of 427–428 kJ. Yes, it is different than the value in the data book.

12. Two reactions occurring in the manufacture of sulfuric acid are shown below:

reaction I $S(s) + O_2(g) \rightarrow SO_2(g)$ $\Delta H^\ominus = -297$ kJ

reaction II $SO_2(g) + \frac{1}{2}O_2(g) \rightleftharpoons SO_3(g)$ $\Delta H^\ominus = -92$ kJ

(i) State the name of the term $\Delta H^\ominus$. State, with a reason, whether reaction I would be accompanied by a decrease or increase in temperature. **(3)**

ΔH° = standard enthalpy (change) of rxn

Temperature increase

Reaction is exothermic (OR) sign of ΔH° is (−) negative.

(ii) At room temperature sulfur trioxide, SO_3, is a solid. Deduce, with a reason, whether the $\Delta H^\ominus$ value would be more negative or less negative if $SO_3(s)$ instead of $SO_3(g)$ were formed in reaction II. **(2)**

More negative (exothermic) + b/c heat is given out when gas changes to solid (OR) solid has less enthalpy than gas

(iii) Deduce the $\Delta H^\ominus$ value of this reaction: **(1)**

$$S(s) + 1\tfrac{1}{2}O_2(g) \rightarrow SO_3(g)$$

$S(s) + O_2(g) \rightarrow SO_2(g)$ (−297)

$SO_2(g) + \frac{1}{2}O_2(g) \rightarrow SO_3(g)$ (−92)

$S(s) + 1\frac{1}{2}O_2(g) \rightarrow SO_3(g)$ $\Delta H^\circ_{rxn} = -389$ kJ **(Total 6 marks)**

13. (i) Define the term *average bond enthalpy*. **(3)**

Again?

→ The energy needed to break one bond in a molecule in the (gaseous) state, averaged value of all bonds of those elements in similar compounds.

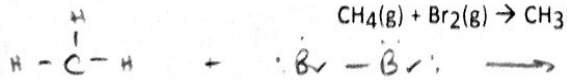
(ii) Explain why Br_2 is not suitable as an example to illustrate the term *average bond enthalpy*. **(1)**

B/c there is no other species w/ Br_2 bonds.

$Br_2 \Rightarrow :\ddot{B}r - \ddot{B}r:$

(iii) Using values from Table 10 of the Data Booklet, calculate the enthalpy change for the following reaction: **(3)**

$$CH_4(g) + Br_2(g) \rightarrow CH_3Br(g) + HBr(g)$$



Broken: $(C-H)\times 4\ (413)$, $(Br-Br)\times 1\ (193)$

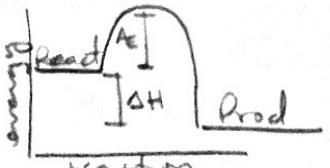
Σ(break) − Σ(make)

1845 kJ − 1895 kJ

$\Delta H^\circ_{rxn} = -50$ kJ

Formed: $(C-H)\times 3\ (413)$, $(C-Br)\times 1\ (290)$, $(H-Br)\times 1\ (366)$

(iv) Sketch an enthalpy level diagram for the reaction in part (iii). **(2)**



(v) Without carrying out a calculation, suggest, with a reason, how the enthalpy change for the following reaction compares with that of the reaction in part (iii). **(2)**

$$CH_3Br(g) + Br_2(g) \rightarrow CH_2Br_2(g) + HBr(g)$$

(Total 11 marks)

This equation should give an answer about the same (if not equal) b/c the same number and types of bonds are being lost & gained.

(overall breaks a C-H and Br-Br, forming both an C-Br and H-Br).